

Entropy Worksheet

Indicate the sign of ΔS for the following processes.

$\text{CaO(s)} + \text{CO}_2\text{(g)} \rightarrow \text{CaCO}_3\text{(s)}$	ΔS -
$\text{HCl(g)} + \text{NH}_3\text{(s)} \rightarrow \text{NH}_4\text{Cl(s)}$	-
$\text{N}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{NO}_2\text{(g)}$	+
$\text{FeCl}_2\text{(s)} + \text{H}_2\text{(g)} \rightarrow \text{Fe(s)} + 2\text{HCl(g)}$	+
The reaction between zinc and hydrochloric acid	+
Silver nitrate combining with potassium bromide	-
Decomposition of lithium chloride	+

Which substance possesses the most entropy?

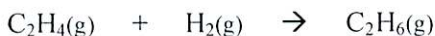
1 mol of $\text{P}_4\text{(g)}$ at 300°C, 0.01 atm or 1 mol of $\text{As}_4\text{(g)}$ at 300°C, 0.01 atm

He(g) at 3 atm pressure or He(g) at 1.5 atm pressure

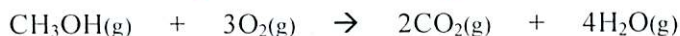
0.5 mol of $\text{N}_2\text{(g)}$ at 298 K and 20 L volume or 0.5 mol $\text{CH}_4\text{(g)}$ at 298 K and 20 L volume

100 g of $\text{Na}_2\text{SO}_4\text{(s)}$ at 30°C or 100 g of $\text{Na}_2\text{SO}_4\text{(aq)}$ at 30°C

Calculate the ΔS° using standard entropy values from the tables in the back of the text. In each case explain the sign of ΔS° .



$$\Delta S^\circ = 229.5 \frac{\text{J}}{\text{K}} - \left(219 \frac{\text{J}}{\text{K}} + 131 \frac{\text{J}}{\text{K}} \right) = -120.5 \frac{\text{J}}{\text{K}} \quad \begin{array}{l} 2 \text{ gas} \rightarrow 1 \text{ gas} \\ \text{makes sense} \end{array}$$



$$\Delta S^\circ = \left(2(214 \frac{\text{J}}{\text{K}}) + 4(189 \frac{\text{J}}{\text{K}}) \right) - \left(240 \frac{\text{J}}{\text{K}} + 3(205 \frac{\text{J}}{\text{K}}) \right) = 329 \frac{\text{J}}{\text{K}} \quad \text{More gas molecules made}$$



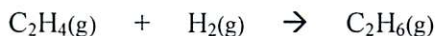
$$\Delta S^\circ = \left(3(189 \frac{\text{J}}{\text{K}}) + 2(28 \frac{\text{J}}{\text{K}}) \right) - \left(3(131 \frac{\text{J}}{\text{K}}) + 51 \frac{\text{J}}{\text{K}} \right) = 179 \frac{\text{J}}{\text{K}} \quad \leftarrow \text{H}_2\text{O has more vibrational motion}$$

Explain why the hydroxide ion has a ΔS° of -10.7 J/(mol·K)

$\text{OH}^- \text{aq}$ b/c the (-) charge make H_2O less Random

Free Energy Worksheet

Calculate the free energy change for the following processes, using the tabulated free energy values at the back of the book. Are these reactions spontaneous?



$$\Delta G^\circ = -32.9 \frac{\text{kJ}}{\text{mol}} - (68 \frac{\text{kJ}}{\text{mol}} - 0) = -100.9 \text{ kJ} \quad \text{yes spontaneous}$$

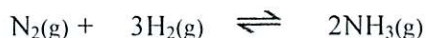


$$\Delta G^\circ = ((2(0) + 3(-229 \text{ kJ})) - ((-1582 \text{ kJ}) + 3(0))) = +895 \text{ kJ} \quad \text{not spontaneous}$$



$$\Delta G^\circ = (2(-95 \text{ kJ}) + -377) - (2(-384 \text{ kJ}) + (-237 \text{ kJ})) = 438 \text{ kJ} \quad \text{Not spontaneous}$$

For the following reaction:



Calculate ΔG at 298 K for a reaction mixture that has 1.0 atm N_2 , 3.0 atm H_2 , and 0.5 atm NH_3 .

$$\Delta G^\circ = 2(-17 \text{ kJ}) - (0 + 3(0)) = -34 \text{ kJ}$$

$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$-34 \text{ kJ} + \left(\frac{1 \text{ kJ}}{1000 \text{ J}}\right) \left(8.315 \frac{\text{J}}{\text{mol K}}\right) (298 \text{ K}) \left(\ln \frac{(0.5 \text{ atm})^2}{(1.0 \text{ atm})(3.0 \text{ atm})^3}\right) = -46 \text{ kJ}$$

Calculate the equilibrium constant for the reaction in the previous question.

$$\Delta G^\circ = -RT \ln K$$

$$-34 \text{ kJ} = -\left(\frac{1 \text{ kJ}}{1000 \text{ J}}\right) (8.315 \frac{\text{J}}{\text{mol K}}) (298 \text{ K}) (\ln K)$$

$$13.7 = \ln K$$

$$K = 9.1 \times 10^5$$

Using the K_a for nitrous acid, calculate the ΔG° for the dissociation of this acid. What is the value of ΔG for the reaction at equilibrium? What is the value of ΔG when $[\text{H}^+]$ is $5.0 \times 10^{-2} \text{ M}$, $[\text{NO}_2^-]$ is $6.0 \times 10^{-4} \text{ M}$, and $[\text{HNO}_2]$ is 0.20 M .

$$K_a = 4.0 \times 10^{-4}$$

$$\Delta G^\circ = -RT \ln K$$

$$\Delta G^\circ = -\left(8.315 \frac{\text{J}}{\text{mol K}}\right) (298 \text{ K}) (\ln 4.0 \times 10^{-4}) \left(\frac{1 \text{ kJ}}{1000 \text{ J}}\right)$$

$$= 19.4$$

$$\Delta G_{\text{eq}} = 0$$

$$\Delta G = \Delta G^\circ + RT \ln Q$$

$$19.4 \text{ kJ} + \frac{1 \text{ kJ}}{1000 \text{ J}} \left(8.315 \frac{\text{J}}{\text{mol K}}\right) (298 \text{ K}) \left(\ln \frac{(5.0 \times 10^{-2})(6.0 \times 10^{-4})}{0.20}\right)$$

$$= -2.12 \text{ kJ}$$